

Problems - Chapter 13 (with solutions)

1) Define the following terms: a) solution, b) miscible, c) hydration, d) percent by mass (solute), e) colligative property, f) hypotonic.

a) solution - A homogeneous mixture, that is, a mixture whose composition is the same throughout.

b) miscible - Refers to two liquids that will always form a solution no matter how much of each liquid is used.

c) hydration - The process of surrounding solute particles in aqueous solution with water molecules. Hydration of ions occurs by ion-dipole attractive forces, and is particularly strong.

$$\text{d) percent by mass (solute)} = \frac{\text{mass solute}}{\text{mass solution}} \times 100 \%$$

e) colligative property - A property of a solution of a nonvolatile solute and a volatile solvent which at most (for ideal conditions) depends only on the physical properties of the solvent and the concentration of solute particles.

f) hypotonic - A solution with a lower concentration of dissolved substances than blood plasma. Red blood cells will swell, and can burst open, if placed in a hypotonic solution.

2) The following question concerns mixing of liquids.

a) Consider mixing two liquids where mixing is exothermic ($\Delta H_{\text{soln}} < 0$). Would you expect a solution to form (yes/maybe/no)? Justify your answer.

Yes. Mixing increases randomness (entropy), which favors solution formation. Since the mixing is exothermic, mixing lowers the energy, which also favors solution formation. Since both factors favor solution formation, I would expect a solution to form.

b) Consider mixing two liquids where mixing is endothermic ($\Delta H_{\text{soln}} > 0$). Would you expect a solution to form (yes/maybe/no)? Justify your answer.

Maybe. Mixing increases randomness (entropy), which favors solution formation. However, the mixing is endothermic, and so mixing raises the energy, which favors not forming a solution. Since one factor favors solution formation and one factor does not, a solution might form. Whether a solution forms depends on which of the two opposing factors is more important, and so will mainly depend on how large the value is for ΔH_{soln} . The larger the value for ΔH_{soln} (the more endothermic the mixing process) the less likely a solution will form.

3) (13.12) Explain the variation in the solubility in water of the listed alcohols:

<u>Compound</u>	<u>Solubility (g/100 g H₂O) at T = 20. °C</u>
CH ₃ OH	∞
CH ₃ CH ₂ OH	∞
CH ₃ CH ₂ CH ₂ OH	∞
CH ₃ CH ₂ CH ₂ CH ₂ OH	9.0
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH	2.7

∞ means the alcohol and water are completely miscible

The -OH group of the alcohols can form strong hydrogen bonds with water molecules, but this property diminishes in importance as the chain length increases.

The longer the C-C chain, the less significant the -OH bond, and the more the molecule "looks like" a nonpolar hydrocarbon. Hence, as the C-C chain length increases, the molecule becomes less polar. Since "like dissolves like", as the molecules become less polar, the solubility in water (a polar solvent) decreases.

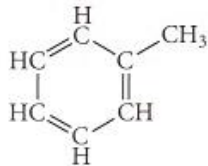
4) What does the statement like dissolves like mean with respect to solution formation?

The statement "like dissolves like" means that polar solutes tend to dissolve well in polar solvents, and that nonpolar solutes tend to dissolve well in nonpolar solvents. Polar solutes usually do not dissolve well in nonpolar solvents, and nonpolar solutes usually do not dissolve well in polar solvents.

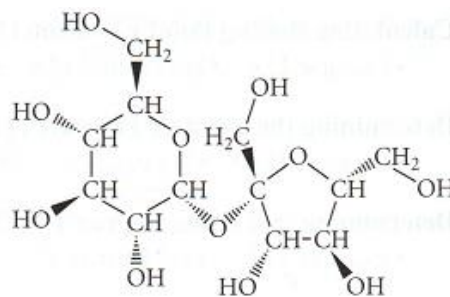
The underlying explanation for this "rule" is that $\Delta H_{\text{soln}} \cong 0$ when the solute and the solvent are both polar or when they are both nonpolar. Solution formation will usually occur in this case because of the increase in randomness (entropy) when a solution forms. When the solvent is polar and the solute is nonpolar, or when the solvent is nonpolar and the solute is polar, it is usually the case that $\Delta H_{\text{soln}} \gg 0$. The increase in enthalpy (essentially an increase in energy) required to form a solution, generally prevents solution formation from taking place.

5) For each compound would you expect greater solubility in water (H₂O) or in hexane (C₆H₁₄)? Indicate the kinds of intermolecular forces that would occur between the solute and the solvent in which the molecule is more soluble.

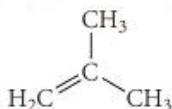
a. toluene



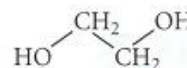
b. sucrose
(table sugar)



c. isobutene



d. ethylene glycol



Molecules a and c (toluene and isobutene) are both hydrocarbons (they contain only carbon and hydrogen atoms). The molecules are nonpolar. Therefore these two substances should dissolve well in hexane (a nonpolar solvent) but not dissolve well in water (a polar solvent). They will interact with the hexane by dispersion forces.

Molecules b and d (sucrose and ethylene glycol) contain several polar –OH bonds, and so the molecules are polar. Therefore these two molecules should dissolve well in water (a polar solvent) but not dissolve well in hexane (a nonpolar solvent). They will interact with water by hydrogen bonding (a type of dipole-dipole interaction) and dispersion forces.

6) (13.18) Calculate the molality of each of the following aqueous solutions.

a) 2.55 M NaCl solution, density = $D = 1.08 \text{ g/mL}$.

Recall that molality = $\frac{\text{moles solute}}{\text{kg solvent}}$

$MW(\text{NaCl}) = 58.44 \text{ g/mol}$

Assume 1000.0 mL of solution. (Note that you will get the same answer if you assume a different volume of solution, but 1000.0 mL is a convenient amount to work with.)

Then mass solution = $1000.0 \text{ mL soln} \frac{1.08 \text{ g soln}}{\text{mL soln}} = 1080. \text{ g soln} = 1.080 \text{ kg soln}$

moles NaCl = $1.000 \text{ L soln} \frac{2.55 \text{ mol NaCl}}{\text{L soln}} = 2.55 \text{ mol NaCl}$

mass of water = mass soln - mass NaCl

$$= 1080. \text{ g} - [2.55 \text{ mol NaCl} \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}}] = 931. \text{ g} = 0.931 \text{ kg}$$

And so molality = $\frac{2.55 \text{ mol NaCl}}{0.931 \text{ kg water}} = 2.74 \text{ mol/kg}$

b) 45.2 percent by mass KBr.

Here it is convenient to assume we have 100.0 g solution. Since the solution is 45.2 % by mass KBr, 100 g of the solution contains 45.2 g KBr and 54.8 g H₂O. Since MW(KBr) = 119.0 g/mol, then

$$\text{moles KBr} = 45.2 \text{ g KBr} \frac{1 \text{ mol KBr}}{119.0 \text{ g KBr}} = 0.380 \text{ mol KBr}$$

$$\text{kg water} = 54.8 \text{ g water} \frac{1 \text{ kg}}{1000. \text{ g}} = 0.0548 \text{ kg water}$$

And so molality = $\frac{0.380 \text{ mol KBr}}{0.0548 \text{ kg water}} = 6.93 \text{ mol/kg}$

7) (13.24) The density of an aqueous solution containing 15.0 percent ethyl alcohol (EtOH = CH₃CH₂OH) by mass is D = 0.984 g/mL.

a) Calculate the molality of this solution.

Assume 100.0 g of solution. (Note that you will get the same answer if you assume a different mass of solution, but 100.0 g is an easy amount to work with.)

The mass of ethyl alcohol in 100.0 g solution is 0.150 (100.0 g) = 15.0 g EtOH

The mass of water in 100.0 g solution is 100.0 g - 15.0 g = 85.0 g = 0.0850 kg water

The molecular mass of ethyl alcohol is MW = 46.07 g/mol, and so the moles of ethyl alcohol in 100.0 g solution is

$$\text{moles EtOH} = 15.0 \text{ g EtOH} \frac{1 \text{ mol}}{46.07 \text{ g}} = 0.3256 \text{ mol EtOH}$$

The molality of ethyl alcohol is then $m = \frac{0.3256 \text{ mol EtOH}}{0.0850 \text{ kg water}} = 3.83 \text{ mol/kg}$

b) Calculate the molarity of the solution

The volume occupied by 100.0 g of solution is

$$V = 100.0 \text{ g soln} \frac{1 \text{ mL}}{0.984 \text{ g}} = 102. \text{ mL} = 0.102 \text{ L}$$

The moles of EtOH is 0.3256 mol (from part a). Therefore

$$\text{molarity} = \frac{0.3256 \text{ mol EtOH}}{0.102 \text{ L soln}} = 3.19 \text{ mol/L}$$

c) What volume of solution would contain 0.250 moles of ethyl alcohol?

$$V = 0.250 \text{ mol EtOH} \frac{1 \text{ L}}{3.19 \text{ mol}} = 0.0784 \text{ L} = 78.4 \text{ mL soln}$$

8) An aqueous KNO_3 solution is made using 72.5 g of KNO_3 dissolved to a total solution volume of 2.00 L. Calculate the molarity, molality, and mass percent KNO_3 in the solution. Assume a density of 1.050 g/mL for the solution.

$$\text{MW}(\text{KNO}_3) = 101.1 \text{ g/mol} \quad \text{MW}(\text{H}_2\text{O}) = 18.02 \text{ g/mol}$$

$$\text{g}(\text{KNO}_3) = 72.5 \text{ g} \quad \text{mol}(\text{KNO}_3) = 72.5 \text{ g} \frac{1 \text{ mol}}{101.1 \text{ g}} = 0.7171 \text{ mol}$$

$$\text{g solution} = 2000 \text{ mL} \frac{1.05 \text{ g}}{1 \text{ mL}} = 2100. \text{ g}$$

Since $\text{g solution} = \text{g water} + \text{g KNO}_3$,

$$\text{g water} = \text{g solution} - \text{g KNO}_3 = 2100. \text{ g} - 72.5 \text{ g} = 2027.5 \text{ g}$$

So

$$\text{molarity (M)} = \frac{0.7171 \text{ mol KNO}_3}{2.00 \text{ L soln}} = 0.359 \text{ mol/L}$$

$$\text{molality (m)} = \frac{0.7171 \text{ mol KNO}_3}{2.0275 \text{ kg water}} = 0.354 \text{ mol/kg}$$

$$\text{mass percent KNO}_3 = \frac{72.5 \text{ g KNO}_3}{2100. \text{ g soln}} \cdot 100\% = 3.45 \% \text{ KNO}_3 \text{ by mass}$$

9) A stock solution is prepared by dissolving 18.21 g of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) in water. The final volume of the solution is $V = 250.0 \text{ mL}$.

a) What is the molarity of sodium acetate in the solution?

$$\text{MW}(\text{sodium acetate}) = 82.03 \text{ g/mol}$$

$$\text{moles sodium acetate} = 18.21 \text{ g} \frac{1 \text{ mol}}{82.03 \text{ g}} = 0.2220 \text{ mol sodium acetate}$$

$$\text{molarity} = \frac{0.2220 \text{ mol}}{0.2500 \text{ L}} = 0.8880 \text{ mol/L}$$

b) How many mL of the stock solution of sodium acetate are needed to prepare 100.0 mL of a 0.1000 mol/L sodium acetate solution?

$$M_i V_i = M_f V_f$$

$$\text{and so } V_i = \frac{M_f}{M_i} V_f = \frac{(0.1000 \text{ M})}{(0.8880 \text{ M})} (100.0 \text{ mL}) = 11.36 \text{ mL stock solution}$$

10) The Henry's law constant for methyl bromide (CH_3Br) in water is $0.159 \text{ mol/L}\cdot\text{atm}$ at $T = 25.^\circ\text{C}$. How many grams of methyl bromide will dissolve in 1.000 L of water at this temperature when the pressure of methyl bromide above the water is $p = 320. \text{ torr}$?

Let MeBr = methyl bromide

$$\text{From Henry's law } [\text{MBr}] = k p_{\text{MBr}}$$

$$p_{\text{MeBr}} = 320. \text{ torr} \frac{1 \text{ atm}}{760 \text{ torr}} = 0.421 \text{ atm}$$

$$[\text{MeBr}] = (0.159 \text{ mol/L}\cdot\text{atm}) (0.421 \text{ atm}) = 0.0669 \text{ mol/L}$$

So 1.000 L of water contains 0.0669 moles of methyl bromide. Since $\text{MW}(\text{MeBr}) = 94.9 \text{ g/mol}$

$$\text{grams MeBr} = (0.0669 \text{ mol}) (94.9 \text{ g/mol}) = 6.36 \text{ g}$$

11) A solution contains a mixture of pentane and hexane at room temperature. The solution has a vapor pressure of 258. torr. Pure pentane and hexane have vapor pressures of 425 torr and 151 torr, respectively, at room temperature. What is the mole fraction composition of the mixture? Assume ideal behavior.

Let P = pentane and H = hexane, and assume Raoult's law applies.

$$p_{\text{total}} = p_{\text{P}} + p_{\text{H}} = X_{\text{P}} p_{\text{P}}^{\circ} + X_{\text{H}} p_{\text{H}}^{\circ}$$

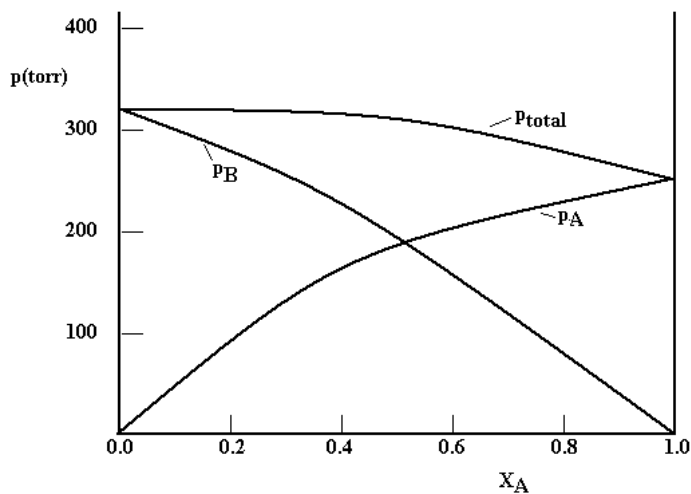
We have one equation and two unknowns. But $X_{\text{P}} + X_{\text{H}} = 1$, so $X_{\text{P}} = 1 - X_{\text{H}}$, so

$$\begin{aligned} p_{\text{total}} &= (1 - X_{\text{H}}) p_{\text{P}}^{\circ} + X_{\text{H}} p_{\text{H}}^{\circ} = p_{\text{P}}^{\circ} - X_{\text{H}} p_{\text{P}}^{\circ} + X_{\text{H}} p_{\text{H}}^{\circ} \\ &= p_{\text{P}}^{\circ} + X_{\text{H}} (p_{\text{H}}^{\circ} - p_{\text{P}}^{\circ}) \end{aligned}$$

$$\text{Or, finally, } X_{\text{H}} = \frac{(p_{\text{total}} - p_{\text{P}}^{\circ})}{(p_{\text{H}}^{\circ} - p_{\text{P}}^{\circ})} = \frac{(258 - 425) \text{ torr}}{(151 - 425) \text{ torr}} = 0.609$$

$$\text{And so } X_{\text{P}} = 1 - 0.609 = 0.391$$

12) The diagram below is for a solution of two volatile liquids A and B, and shows the partial pressure of A, the partial pressure of B, and the total pressure above a solution as a function of X_{A} , the mole fraction of A in the solution.



a) Do A and B form an ideal solution? Justify your answer.

No, A and B do not form an ideal solution. For an ideal solution the partial pressures of A and B are given by Raoult's law, and so a plot of partial pressure or total pressure vs X_{A} should give straight lines. Since that is not observed here, A and B do not obey Raoult's law and so do not form an ideal solution.

b) If A and B do not form an ideal solution, are the A--B interactions stronger than, approximately equal to, or weaker than the average of the A---A and B---B interactions? Justify your answer.

The partial pressures of A and B are higher than we would get if Raoult's law were obeyed. That means that the molecules of A and B are less strongly held in the liquid than if Raoult's law were obeyed. Therefore, the A---B interactions are weaker than the average of the A---A and B---B interactions.

13) How many grams of glucose ($C_6H_{12}O_6$) must be added to 552. g of water to give a solution with a vapor pressure 0.20 torr less than that of pure water at 20.0 °C? The vapor pressure of pure water at this temperature is $p^{\circ}_{H_2O} = 17.5$ torr.

For vapor pressure lowering $\Delta p = X_B p^{\circ}$

$$\text{So } X_{\text{glu}} = \frac{\Delta p}{p^{\circ}} = \frac{0.20 \text{ torr}}{17.5 \text{ torr}} = 0.0114$$

$$\text{And so } X_{H_2O} = 1.000 - 0.0114 = 0.9886$$

The number of moles of H_2O (MW = 18.02 g/mol) is

$$n(H_2O) = \frac{552. \text{ g}}{(18.02 \text{ g/mol})} = 30.63 \text{ mol } H_2O$$

$$\text{Since } \frac{n(\text{glucose})}{n(H_2O)} = \frac{X_{\text{glu}}}{X_{H_2O}}$$

$$n(\text{glucose}) = n(H_2O) \frac{X_{\text{glu}}}{X_{H_2O}} = (30.63 \text{ mol}) \frac{0.0114}{0.9886} = 0.353 \text{ mol glucose}$$

Since MW(glucose) = 180.2 g/mol

$$g(\text{glucose}) = 0.353 \text{ mol } \frac{180.2 \text{ g}}{1 \text{ mol}} = 63.6 \text{ g glucose}$$

14) An ethylene glycol solution contains 21.2 g of ethylene glycol ($C_2H_6O_2$) in 85.4 mL of water. Determine the freezing point and boiling point of the solution. Assume a density of 1.000 g/mL for water. Note that for water $K_f = 1.86$ °C•kg/mol and $K_b = 0.512$ °C•kg/mol, and that ethylene glycol is a nonvolatile and nonionizing solute.

To use the equations for boiling point elevation and freezing point depression we need to find the molality of ethylene glycol in the solution. Let ET = ethylene glycol.

$$\text{MW(ET)} = 62.07 \text{ g/mol}$$

So the moles of ethylene glycol is

$$n_{\text{ET}} = \frac{21.2 \text{ g}}{62.07 \text{ g/mol}} = 0.3415 \text{ mol ET}$$

There are 85.4 mL of water, and since the density of water is 1.00 g/mL, there are 85.4 g (or 0.0854 kg) of water. So the molality of the solution is

$$m = \frac{0.3415 \text{ mol ET}}{0.0854 \text{ kg water}} = 4.00 \text{ mol/kg}$$

The formulas for boiling point elevation and freezing point depression are

$$\Delta T_b = K_b m \quad K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

$$\Delta T_f = K_f m \quad K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

$$\text{So } \Delta T_b = K_b m = (0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}) (4.00 \text{ mol/kg}) = 2.05 \text{ }^\circ\text{C}$$

$$\Delta T_f = K_f m = (1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}) (4.00 \text{ mol/kg}) = 7.44 \text{ }^\circ\text{C}$$

Since the normal boiling point and normal freezing point for water are 100.00 °C and 0.00 °C, respectively, the boiling point and freezing point for the solution are

$$T_b = 100.00 \text{ }^\circ\text{C} + 2.05 \text{ }^\circ\text{C} = 102.05 \text{ }^\circ\text{C}$$

$$T_f = 0.00 \text{ }^\circ\text{C} - 7.44 \text{ }^\circ\text{C} = -7.44 \text{ }^\circ\text{C}$$

Note that the boiling point of the solution is higher than that for pure water and the freezing point of the solution is lower than that for pure water.

15) An aqueous solution containing 35.9 g of an unknown nonvolatile and nonionizing molecular compound in 150.0 g of water was found to have a freezing point of $-1.30 \text{ }^\circ\text{C}$. Calculate the molar mass of the unknown compound. For water, $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$.

For freezing point depression $\Delta T_f = K_f m$, where m = molality of solute particles

$$\text{So } m = \frac{\Delta T_f}{K_f} \quad \Delta T_f = T_f^\circ - T_f = 0.0^\circ\text{C} - (-1.3 \text{ }^\circ\text{C}) = 1.3 \text{ }^\circ\text{C}$$

$$\text{So } m = \frac{1.3 \text{ }^\circ\text{C}}{(1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol})} = 0.699 \text{ mol/kg}$$

There are 150.0 g of solvent (water). The number of moles of solute is

$$n_{\text{solute}} = 0.1500 \text{ kg solvent} \frac{0.699 \text{ mol}}{1 \text{ kg}} = 0.105 \text{ mol solute}$$

$$\text{So MW} = \frac{\text{g solute}}{\text{mol solute}} = \frac{35.9 \text{ g}}{0.105 \text{ mol}} = 342. \text{ g/mol}$$

16) (13.66) What is the osmotic pressure, in atmospheres, of a 1.57 M aqueous solution of urea ((NH₂)₂CO), a nonvolatile and nonionizing solute, at T = 27. °C.

$$\Pi = MRT = (1.57 \text{ mol/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(27. + 273.) \text{ K} = 38.7 \text{ atm}$$

17) (13.78) Arrange the following aqueous solutions in order of decreasing freezing point, and explain your reasoning: 0.50 m HCl, 0.50 m glucose, 0.50 m CH₃COOH.

The equation for freezing point depression, including the van't Hoff factor (i), is

$$\Delta T_f = i K_f m$$

The freezing point will be depressed most by the solution that contains the most solute particles. Since all of the solutions have the same concentration of solute, the freezing point will be depressed most by the solute with the largest value for I, the van't Hoff factor.

We can classify each solute as a strong electrolyte, a weak electrolyte, or a nonelectrolyte. HCl is a strong electrolyte, so under ideal conditions it will completely dissociate into two particles per molecule (i = 2). The concentration of particles will be 1.00 m. Acetic acid is a weak electrolyte, so it will only dissociate to a small extent (1 < i < 2). The concentration of particles will be greater than 0.50 m but less than 1.00 m. Glucose is a nonelectrolyte (i = 1), so glucose molecules remain as glucose molecules in solution. The concentration of particles will be 0.50 m. For these solutions, the order of freezing points is:

$$0.50 \text{ m glucose} > 0.50 \text{ m acetic acid} > 0.50 \text{ m HCl}$$

The HCl solution will have the lowest freezing point (greatest freezing point depression), and the glucose solution will have the highest freezing point (smallest freezing point depression).

18) (13.86) A quantity of 7.480 g of an organic compound is dissolved in water to make 300.0 mL of solution. The solution has an osmotic pressure of 1.43 atm at $T = 27.^\circ\text{C}$. The analysis of this compound shows that it contains 41.8 percent C, 4.7 percent H, 37.3 percent O, and 16.3 percent N. Find the molecular formula for the compound.

Find the concentration of the solution, then work out the molar mass. Note that $27.^\circ\text{C} = 300. \text{K}$. The concentration is:

$$M = \frac{\Pi}{RT} = \frac{(1.43 \text{ atm})}{(0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(300. \text{K})} = 0.0581 \text{ mol/L}$$

The solution volume is 0.3000 L so the number of moles of solute is:

$$n = 0.3000 \text{ L} \frac{0.0581 \text{ mol}}{\text{L}} = 0.0174 \text{ mol}$$

The molar mass is then

$$\text{MW} = \frac{7.480 \text{ g}}{0.0174 \text{ mol}} = 430. \text{ g/mol}$$

The empirical formula can be found most easily by assuming a 100.0 g sample of the substance.

$$\text{moles C} = 41.8 \text{ g} \frac{1 \text{ mol}}{12.01 \text{ g}} = 3.48 \text{ mol C} / 1.16 = 3.00$$

$$\text{moles H} = 4.7 \text{ g} \frac{1 \text{ mol}}{1.008 \text{ g}} = 4.7 \text{ mol H} / 1.16 = 4.05$$

$$\text{moles O} = 37.3 \text{ g} \frac{1 \text{ mol}}{16.00 \text{ g}} = 2.33 \text{ mol O} / 1.16 = 2.01$$

$$\text{moles N} = 16.3 \text{ g} \frac{1 \text{ mol}}{14.01 \text{ g}} = 1.16 \text{ mol N} / 1.16 = 1.00$$

This gives the empirical formula, $\text{C}_3\text{H}_4\text{O}_2\text{N}$, which has a mass of 86.0 g per formula unit. The molar mass is five times this amount ($430 \div 86.0 = 5.0$), so the molecular formula is $(\text{C}_3\text{H}_4\text{O}_2\text{N})_5$ or $\text{C}_{15}\text{H}_{20}\text{O}_{10}\text{N}_5$.

19) A solution is prepared by dissolving 2.18 g of potassium nitrate (KNO_3 , MW = 101.1 g/mol) in water. The final volume of the solution is $V = 400.0 \text{ mL}$. What is the osmotic pressure of the solution at $T = 20.0 \text{ }^\circ\text{C}$?

The number of moles of KNO_3 is

$$n = 2.18 \text{ g} \frac{1 \text{ mol}}{101.1 \text{ g}} = 0.0216 \text{ mol}$$

and so the molarity of the solution is $M = \frac{0.0216 \text{ mol } \text{KNO}_3}{0.4000 \text{ L}} = 0.0539 \text{ mol/L } \text{KNO}_3$

However, KNO_3 is an ionic compound $\text{KNO}_3(\text{s}) \rightarrow \text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

and so the concentration of solute particles is

$$\frac{0.0539 \text{ mol } \text{KNO}_3}{\text{L soln}} \frac{2 \text{ mol particle}}{1 \text{ mol } \text{KNO}_3} = 0.1078 \text{ mol particles}$$

The osmotic pressure is therefore

$$\Pi = [\text{B}]RT = (0.1078 \text{ mol/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(293.1 \text{ K}) = 2.59 \text{ atm}$$

20) Consider 0.0100 mol/kg aqueous solutions of iron III chloride (FeCl_3), calcium nitrate ($\text{Ca}(\text{NO}_3)_2$), and copper II sulfate (CuSO_4). Which of these solutions will have the highest freezing point, and which of these solutions will have the highest melting point? Justify your answer.

The formulas for boiling point elevation and freezing point depression

$$\Delta T_b = K_b m_B$$

$$\Delta T_f = K_f m_B$$

The solutions all have the same molality. Since boiling point increases, the compound forming the largest number of particles per formula unit will have the highest boiling point. Since freezing point decreases (and recall that freezing point and melting point occur at the same temperature, and are just inverse processes), the compound forming the smallest number of particles per formula unit will have the highest melting point (think about why this is true).

FeCl_3 – 4 particles per formula unit

$\text{Ca}(\text{NO}_3)_2$ – 3 particles per formula unit

CuSO_4 – 2 particles per formula unit

So the FeCl_3 solution will have the highest boiling point, and the CuSO_4 solution will have the highest melting point.